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CHEMISTRY

REVISION NOTES

REDOX

Redox

Oxidation:

- Oxidation is the process in which an atom, ion, or molecule loses electrons.
- It results in an increase in the oxidation state of the species involved.
- Oxidation often involves the addition of oxygen or loss of hydrogen.

Reduction:

- Reduction is the process in which an atom, ion, or molecule gains electrons.
- It results in a decrease in the oxidation state of the species involved.
- Reduction often involves the addition of hydrogen or loss of oxygen.

Oxidizing Agent

An oxidising agent is a substance that accepts electrons from another substance during a chemical reaction.

Reducing Agent

Now, this one is the electron donor. A reducing agent is a substance that gives away electrons to another substance during a chemical reaction.

So, in terms of electrons:

- **Oxidising Agent:** Gains electrons (by making others lose electrons).
- **Reducing Agent:** Loses electrons (by giving them away to others)

Rules for Assigning Oxidation States:

1. **Elements in their natural state:** In their elemental form, atoms have an oxidation state of 0. For example, O_2, H_2, N_2, Cl_2 all have oxidation states of 0.

2. **Monoatomic ions:** The oxidation state of a monatomic ion is equal to its charge. For example, the oxidation state of Na^+ is +1, and the oxidation state of Cl^- is -1.

3. **Hydrogen:** Hydrogen typically has an oxidation state of +1 when combined with nonmetals and -1 when combined with metals.

4. **Oxygen:** Oxygen typically has an oxidation state of -2 in compounds. There are some exceptions, such as in peroxides (e.g., H_2O_2), where its oxidation state is -1.

5. **Alkali metals and alkaline earth metals:** Alkali metals (e.g., Li, Na, K) have an oxidation state of +1, and alkaline earth metals (e.g., Mg, Ca) have an oxidation state of +2.

6. **Fluorine:** Fluorine always has an oxidation state of -1 in compounds.

7. **The sum of oxidation states:** In a neutral compound, the sum of the oxidation states of all atoms must equal zero. In a polyatomic ion, the sum of oxidation states should equal the charge of the ion.

8. **Oxidation states in complex ions:** In complex ions, consider the charge on the ion as a whole when determining the oxidation state of each element within the ion.

9. **Redox reactions:** In a redox reaction, the substance that is oxidised has its oxidation state increased, while the substance that is reduced has its oxidation state decreased.

10. **Change in oxidation state:** Be aware of the change in oxidation state for each element in a chemical reaction. The change in oxidation state is equal to the number of electrons transferred.

Redox Equation (Reduction-Oxidation Equation):

A redox equation represents a chemical reaction in which there is a transfer of electrons from one substance (reductant or reducing agent) to another substance (oxidant or oxidizing agent).

In a redox equation, you typically have two half-reactions: one representing the oxidation half (loss of electrons), and the other representing the reduction half (gain of electrons).

These half-reactions are balanced so that the number of electrons lost in the oxidation half is equal to the number of electrons gained in the reduction half.

For example:

Half-Reaction 1 (Oxidation):

$Cu \rightarrow Cu^{2+} + 2e^-$

Half-Reaction 2 (Reduction):

$Ag^+ + e^- \rightarrow Ag$

Overall Redox Equation:

$Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$

This is a simple example of a redox equation involving the oxidation of copper (Cu) and the reduction of silver ions (Ag^+).

Balancing more complex equations.

Balanced Reduction Half-Equation for NO_3^- to NO_2^-

1. Write the initial half-equation without balancing electrons

$NO_3^- \rightarrow NO_2^-$

2. Balance the nitrogen atoms by adding a coefficient of 1 to NO3- and NO2-:

NO_3- → NO_2-

3. Balance the oxygen atoms by adding water (H2O) to the product side:

NO_3- → NO_2- + H_2O

4. Balance the hydrogen atoms by adding hydrogen ions (H+) to the reactant side:

NO_3- + H+ → NO_2- + H_2O

5. Now, add electrons (e-) to balance the change in oxidation number, which is 2 electrons in this case.

   NO_3- + H+ + 2e- → NO_2- + H_2O

   This balanced half-equation represents the reduction of nitrate ion (NO_3-) to nitrite ion (NO_2-) by first balancing the atoms and then adding 2 electrons to balance the change in oxidation number.

<u>Oxidation Half-Equation for Zn to Zn^{2+}</u>

1. Identify the oxidation numbers and changes:

   Zinc (Zn) goes from 0 to +2 (an oxidation).

2. Write the initial half-equation without balancing electrons:

   $Zn \rightarrow Zn^{2+}$

3. Add electrons (e-) to balance the change in oxidation number:

   $Zn \rightarrow Zn^{2+}$ + 2e-

4. This half-equation represents the oxidation of zinc (Zn) to form zinc ions (Zn^2+) with the loss of 2 electrons

<u>Combine oxidation and reduction half-equations for zinc (Zn) and nitrate (NO^{3-}) to nitrite (NO^{2-}) into a complete redox equation.</u>

Here are the balanced half-equations for reference:

Oxidation Half-Equation for Zn to $Zn^{2+:}$

- $Zn \rightarrow Zn^{2+}$ + 2e-

Reduction Half-Equation for NO^{3-} to NO2-:

NO^{3-} + H+ + 2e- → NO^{2}- + H_2O

To combine them, you must ensure that the number of electrons lost in the oxidation half-equation matches the number of electrons gained in the reduction half-equation.

You can do this by multiplying one or both of the equations to ensure the electrons are equal.

Since the oxidation half-equation involves the loss of 2 electrons, and the reduction half-equation gains 2 electrons, you can combine them directly without any need to multiply:

$Zn + NO^{3-} + H+ \rightarrow Zn^{2+} + NO_2- + H_2O$

This is the balanced redox equation for the reaction in which zinc (Zn) is oxidised to form zinc ions (Zn^2+), and nitrate ions (NO^{3-}) are reduced to form nitrite ions (NO^{2-}) in the presence of hydrogen ions (H+) and water (H_2O).

Exam Questions

State, in terms of electrons, the meaning of the term oxidising agent.

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$Cr_2O_7^{2-}$ can oxidise SO_3^{2-} in acidic conditions to form Cr^{3+} and SO_4^{2-}

Deduce a half-equation for the oxidation of SO_3^{2-} to SO_4^{2-}

Deduce a half-equation for the reduction of $Cr_2O_7^{2-}$ to Cr^{3+}

Deduce the overall equation for the oxidation of SO_3^{2-} by $Cr_2O_7^{2-}$

Half-equation for the oxidation of SO_3^{2-} to SO_4^{2-}

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Half-equation for the reduction of $Cr_2O_7^{2-}$ to Cr^{3+}

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Overall equation

Which compound contains chlorine in an oxidation state of +1?

A Cl_2O

B $KClO_3$

C ClF_3

D CCl_4

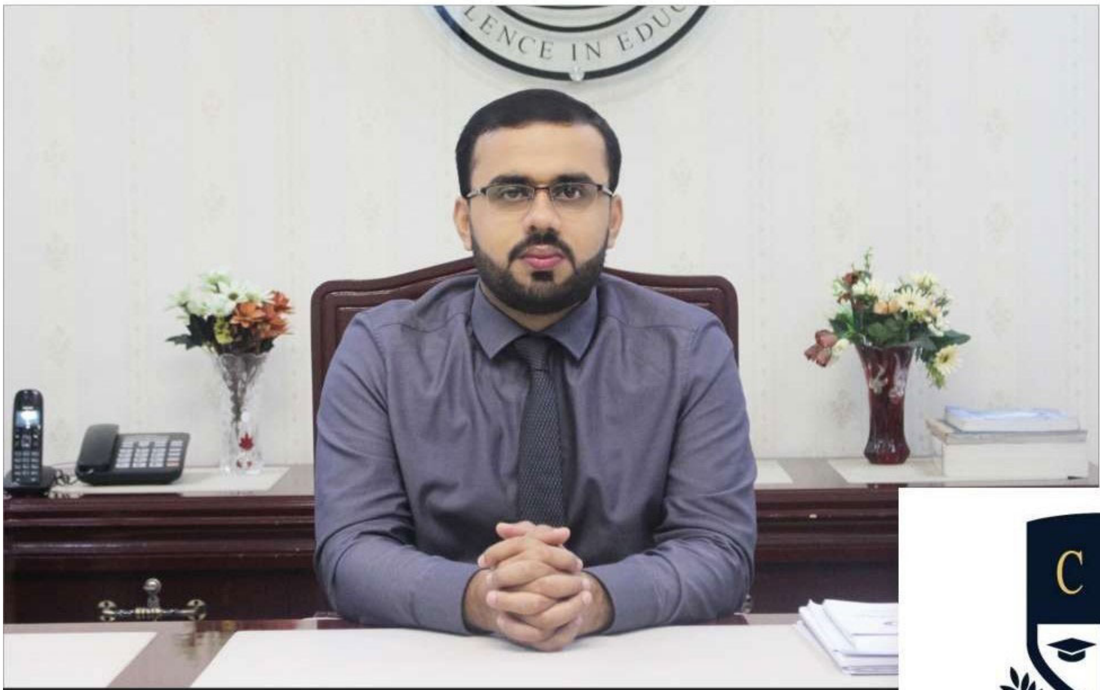
In which conversion is the metal reduced?

A $Cr_2O_7^{2-} \rightarrow CrO_4^{2-}$

B $MnO_4^{2-} \rightarrow MnO_4^{-}$

C $TiO_2 \rightarrow TiO_3^{2-}$

D $VO_3^{-} \rightarrow VO^{2+}$



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